

**Chem 115 POGIL Worksheet - Week #7 - Answers**  
**First Law, Enthalpy, Calorimetry, and Hess's Law**

**Key Questions and Exercises**

1. What are the two major ways in which the internal energy of an object can be categorized? How do these ways differ from one another?

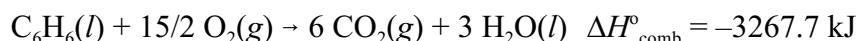
An object can have kinetic energy and potential energy. The kinetic energy arises from motion of the object itself or the fundamental particles of which it is composed. The potential energy arises from its position and composition.

2. If a system loses heat, where does it go?

Because of the First Law of Thermodynamics, heat lost by the system is gained by its surroundings, and vice versa.

3. Describe the following processes as exothermic or endothermic:
- |   |             |
|---|-------------|
| a. an ice cube melts on a warm surface                                    | endothermic |
| b. water freezes  | exothermic  |
| c. calcium chloride is mixed with water, resulting in a very hot solution | exothermic  |
| d. ammonium nitrate is mixed in water, resulting in a very cold solution  | endothermic |

4. The thermochemical equation for the burning of one mole of benzene under standard conditions is



- a. Is this reaction exothermic or endothermic?      exothermic
- b. How much heat is released when a 5.00-g sample of benzene is burned in excess oxygen under standard conditions? (m.w.  $\text{C}_6\text{H}_6 = 78.11 \text{ u}$ )

$$q = (5.00 \text{ g}) \left( \frac{\text{mol}}{78.11 \text{ g}} \right) \left( \frac{-3267.7 \text{ kJ}}{\text{mol}} \right) = -209 \text{ kJ}$$

5. The specific heat of iron (at. wt. = 55.85 u) is 0.450 J/g·K. What is the heat capacity of a 23.5-g block of iron? What is the value of the molar heat capacity of iron?

$$C = (0.450 \text{ J/g}\cdot\text{K})(23.5 \text{ g}) = 10.575 \text{ J/K} = 10.6 \text{ J/K}$$

$$C_m = (0.450 \text{ J/g}\cdot\text{K})(55.85 \text{ g/mol}) = 25.1325 \text{ J/mol}\cdot\text{K} = 25.1 \text{ J/mol}\cdot\text{K}$$

6. How many joules of heat are required to raise the temperature of a 23.5-g block of iron from 25.0 °C to 96.2 °C?

$$\Delta T = (96.2 - 25.0) \text{ }^\circ\text{C} = 71.2 \text{ }^\circ\text{C} = 71.2 \text{ K}$$

$$q = (23.5 \text{ g})(0.450 \text{ J/g}\cdot\text{K})(71.2 \text{ K}) = 752.94 \text{ J} = 753 \text{ J}$$

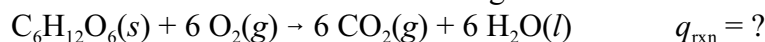
7. A 2.58-g sample of  $\text{KNO}_3$  (f.w. = 101.11 u) was added to 98.57 g of water in a coffee-cup calorimeter. The initial temperature of the water was 22.5 °C, and the temperature of the solution after mixing was 20.4 °C. On the basis of this experiment, what is the heat of solution per mole of  $\text{KNO}_3$  [f.w. = 101.11 u]? The specific heat of water is 4.184 J/g·K.

$$g \text{ soln} = 2.58 \text{ g} + 98.57 \text{ g} = 101.15 \text{ g} \quad \Delta T_{\text{cal}} = 20.4 \text{ }^\circ\text{C} - 22.5 \text{ }^\circ\text{C} = -2.1 \text{ }^\circ\text{C}$$

$$q_{\text{cal}} = (4.184 \text{ J/}^\circ\text{C}\cdot\text{g})(101.15 \text{ g})(-2.1 \text{ }^\circ\text{C}) = (423.21 \text{ J/}^\circ\text{C})(-2.1) = -8.9 \times 10^2 \text{ J}$$

$$\Delta H_{\text{soln}} = \left( \frac{+8.9 \times 10^2 \text{ J}}{2.58 \text{ g KNO}_3} \right) \left( \frac{101.11 \text{ g KNO}_3}{\text{mol KNO}_3} \right) = +3.5 \times 10^4 \text{ J/mol KNO}_3 = +35 \text{ kJ/mol KNO}_3$$

8. The combustion of 1.50 g of glucose,  $\text{C}_6\text{H}_{12}\text{O}_6$  (m.w. = 180.0), caused the temperature of a bomb calorimeter to rise from 19.00 °C to 24.32 °C. The calorimeter had a heat capacity of 4.42 kJ/K. Calculate the heat of combustion of one mole of glucose:



First calculate the heat change for the calorimeter and from it the heat change from burning 1.50 g of glucose.

$$\Delta T = 24.32 \text{ }^\circ\text{C} - 19.00 \text{ }^\circ\text{C} = +5.32 \text{ }^\circ\text{C} = 5.32 \text{ K}$$

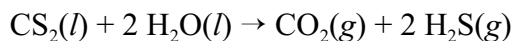
$$q_{\text{cal}} = C_{\text{cal}}\Delta T = (4.42 \text{ kJ/K})(5.32 \text{ K}) = +23.5 \text{ kJ}$$

$$q_{\text{rxn}} = -q_{\text{cal}} = -23.5 \text{ kJ}$$

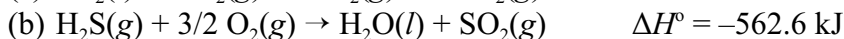
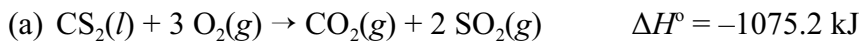
This is the amount of heat liberated by burning 1.50 g of glucose. Now calculate the heat for one mole of glucose.

$$q_{\text{rxn}} = \left( \frac{-23.5 \text{ kJ}}{1.50 \text{ g glucose}} \right) \left( \frac{180.0 \text{ g glucose}}{\text{mol glucose}} \right) = -2820 \text{ kJ/mol} = -2.82 \times 10^3 \text{ kJ/mol}$$

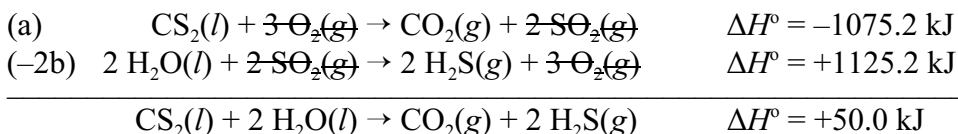
9. Calculate  $\Delta H^\circ$  for the reaction,



Given:

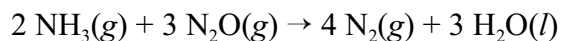


The (a) equation uniquely deals with  $\text{CS}_2(l)$  on the left and  $\text{CO}_2(g)$  on the right, the same places we need them in the sum. Furthermore, the coefficients are the same as what we need in the target equation. Thus, we will use equation (a) once with its  $\Delta H^\circ$  as given. (See first line below.) Equation (b) uniquely deals with  $\text{H}_2\text{S}(g)$  and  $\text{H}_2\text{O}(l)$ , but we need them on the opposite sides and with coefficients of 2 in the target equation. Thus we will reverse equation (b) and multiply it by 2. Its enthalpy contribution to the sum will be  $(-2)(-562.6 \text{ kJ}) = +1125.2 \text{ kJ}$ . (See second line below.) This leads to the following sum:

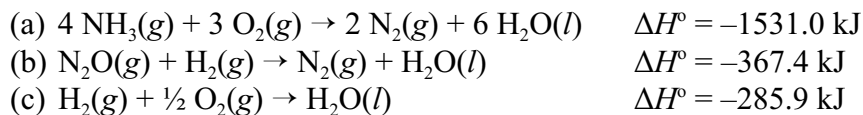


Notice that in analyzing how to use the two given equations, we initially ignored both  $\text{O}_2(g)$  and  $\text{SO}_2(g)$ , because they appeared in more than one equation. Always focus on species that appear uniquely in a given equation, and try to get the correct placement (left or right) and coefficients for those. The species that appear in more than one given equation will add or cancel correctly, assuming that the appropriate set of reference equations is being used.

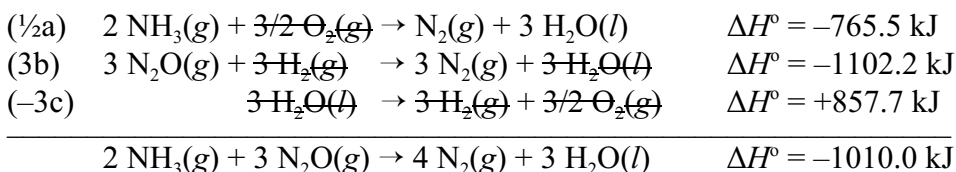
10. Calculate  $\Delta H^\circ$  for the reaction,



Given:



The target equation has 2  $\text{NH}_3(g)$  on the left, and the given (a) equation has 4  $\text{NH}_3(g)$  on the left. Thus, we will divide the (a) equation by 2 and its  $\Delta H^\circ$  value by 2, as well. (See first line below.) The target equation has 3  $\text{N}_2\text{O}(g)$  on the left, and the (b) equation has one  $\text{N}_2\text{O}(g)$  on the left. Thus, we will multiply the (b) equation by 3 and its  $\Delta H^\circ$  value by 3, as well. (See second line below.) The target equation has no  $\text{H}_2(g)$  and no  $\text{O}_2(g)$ , but our sum, so far, has 3  $\text{H}_2(g)$  and  $3/2 \text{O}_2(g)$  on the left. It also at this point has 6  $\text{H}_2\text{O}(l)$  on the right, and we only need 3  $\text{H}_2\text{O}(l)$  in the sum. We need to get rid of the 3  $\text{H}_2(g)$  and  $3/2 \text{O}_2(g)$  on the left and three of the 6  $\text{H}_2\text{O}(l)$  on the right. If we reverse equation (c) and multiply it by 3 we will get all the necessary cancellations. In doing this, we will multiply  $\Delta H^\circ$  for (c) by  $-3$ . The sum, then is

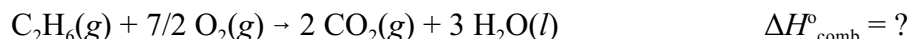


Notice that as we proceeded through the given equations, we did not concern ourselves with the  $\text{N}_2(g)$ , because two of the given equations have it as a product. It would be very difficult to figure out how many times to use the (a) and (b) equations to get the 4  $\text{N}_2(g)$  we need in the target equation. By getting the multiplications right on the basis of  $\text{NH}_3(g)$  from (a) and  $\text{N}_2\text{O}(g)$  from (b), which appear uniquely in those two given equations, we automatically got the correct coefficient for  $\text{N}_2(g)$  in the target equation. Likewise, we did not concern ourselves with getting the 3  $\text{H}_2\text{O}(l)$  coefficient correct from (a) and (b), because both given equations have  $\text{H}_2\text{O}(l)$  in them.

11. The **heat of combustion** for a compound composed of C and H, possibly with either or both O and N, is the heat liberated when *one mole* of the substance is burned with the stoichiometric amount of  $O_2(g)$  to produce,  $CO_2(g)$ ,  $H_2O(l)$ , and  $N_2(g)$ , as may be appropriate.

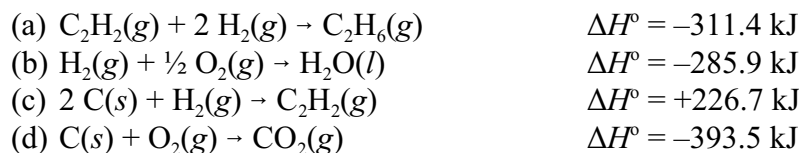
- i. Write the balanced equation for the combustion of one mole of  $C_2H_6(g)$ , which would be the basis for the thermochemical equation defining the heat of combustion of ethane.

The balanced equation that would be the basis for defining the heat of combustion of  $C_2H_6(g)$  is



Note that we have the fractional coefficient 7/2 for  $O_2(g)$  so as to balance this on the basis of one mole of  $C_2H_6(g)$ , because the heat of combustion is defined on the basis of burning one mole of substance. If we balanced this in our usual way, using whole number coefficients, it would be the thermochemical equation for two times the heat of combustion.

- ii. Calculate the heat of combustion of  $C_2H_6(g)$  from the following data:



The target equation we balanced in part i. is the sum of the following given equations, reversed or multiplied as needed:

